Unit 6 – Representing Chemical Change - Objectives

1. Describe chemical changes in terms of rearranging atoms to form new substances.	
2. Recognize that the total number of atoms does not change during a reaction because every reactant atom must be included in a product molecule.	
3. Recognize that the total number of particles (sum of the coefficients) can change during a reaction because of differences in the bonding ratios of each substance	
4. Learn to describe reactions in terms of macroscopic observations	
5. Learn to describe reactions in terms of microscopic behavior of atoms	

6. Learn to write balanced equations to represent these changes symbolically.	
7. Explain that the coefficients in a chemical equation describe the quantities of a. the individual atoms or molecules involved b. the moles of the substances involved.	
8. Observe basic patterns in the way substances react and learn to generalize them to other reactions students encounter. a. Synthesis reactions b. Decomposition reactions c. Combustion reactions d. Single replacement reaction e. Double replacement (ionic) reactions	
9. Describe endo- and exothermic reactions in terms of storage or release of chemical potential energy	

Chemistry - Nail Lab

Purpose

The purpose is to determine the ratio of copper produced to iron consumed in a single replacement reaction.

Procedure

Day 1

1. Label, then mass a 250 mL beaker.

2. Put between 6.0 and 8.5 g of copper(II) chloride in the beaker. To do this, move the rider on the balance beam up by the value you chose, add copper(II) chloride to beaker until the balance is level again.

3. Add about 50 mL distilled water to the beaker. Stir to dissolve the solid.

4. Mass 2 or 3 nails together to ± 0.01g.

5. Place the nails in the copper chloride solution. Observe the reaction; record your observations. Place the labeled beaker in the place designated by your teacher.

Day 2

6. Remove the nails. Rinse or scrape the precipitate (copper metal) from the nails into your labeled 250 mL beaker. Place the nails in a labeled small beaker. Note the appearance of the nails. Place this beaker in the drying oven.

7. Decant solution from the 250 mL beaker. Rinse the precipitate with about 25 mL of distilled water. Try to lose as little of the solid copper as you can when you decant. After a 2^{nd} rinse with distilled water, rinse the copper with 25 mL of 1 M HCl. Rinse one last time with distilled water. Then place the labeled beaker in the drying oven.

8. Mass the dry nails, then discard the nails.

Day 3

9. Mass the beaker + dry copper. Discard the copper in the place designated by your teacher. Wash your beaker and let dry.

Data:

Mass 250 mL beaker	
Mass 250 mL beaker + copper(II) chloride	
Mass nails before reaction	
Mass nails after reaction	
Mass 250 mL beaker + dry copper	

Calculations:

1. Determine the mass of copper produced and the mass of iron used during the reaction.

2. Calculate the moles of copper and moles of iron involved in the reaction.

3. Determine the ratio $\frac{\text{moles of copper.}}{\text{moles of iron}}$

   Express this ratio as an integer. For example, a ratio of 1.33 can be expressed as $\frac{4}{3}$; 0.67 can be expressed as $\frac{2}{3}$, etc.

Conclusion:

1. Why did the reaction stop? Which reactant was used up? How do you know?

2. Describe what was happening to the atoms of iron and copper during the reaction. What is this type of reaction called?

3. What would happen to the ratio of copper to iron if you had placed more nails in the beaker? If you let the reaction go for less time?

4. What is the accepted ratio of copper atoms to iron atoms in this reaction.? Account for differences between your experimental value and the accepted value. Write the balanced equation for the reaction.

Chemistry – Unit 6 Chemical Reactions
Rearranging Atoms

Background

Describe what you already know about each of these ideas. Give an example in each of the last 4 items

Features of Our Current Model of Matter

Conservation of Mass

Chemical Formula

Subscripts in formulas

Coefficient (Hint: what is the function of a coefficient in math?)

Procedure:

1. Use your atom model kit to construct the reactant molecules for each chemical change below. Then rearrange the atoms to form the product molecules. Add more reactant molecules as needed to form complete product molecules with no left-overs.

2. Draw particle diagrams for each reactant molecule used and each product molecule produced under the reaction.

3. Determine the number of each <u>reactant</u> molecule you needed in order to make the product(s) with no leftovers (a complete reaction) and record each number as a coefficient in front of its reactant formula.

4. Determine how many <u>product</u> molecules you would get from the complete reaction. Write that number as a coefficient in front of each product formula.

Name ______________________________

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Rearranging Atoms

Data and Observations:

1. ____H_2 + ____O_2 → ____ H_2O

Diagram:

2. ____H_2 + ____Cl_2 → ____ HCl

Diagram:

3. ____Na + ____O_2 → ____ Na_2O

Diagram:

4. ____N_2 + ____H_2 → ____ NH_3

Diagram:

5. ____CH_4 + ____O_2 → ____CO_2 + ____H_2O

Diagram:

6. ______NO + ____O_2 → ____ NO_2

Diagram:

7. ____Fe + ____Cl_2 → ____ $FeCl_3$

Diagram:

8. ____CH_3OH + ____O_2 → ____CO_2 + ____H_2O

Diagram:

Analysis

1. In each the equation for each reaction, compare the total number of atoms you have before the reaction (reactant atoms) to the total number after the reaction (product atoms).

2. At the beginning of the year we observed that mass is conserved in changes. How does your answer to question 1 explain conservation of mass?

3. Look at the product molecule (ammonia) in reaction #4.
   a. What does the coefficient tell us about this substance?

   b. What do the subscripts on the nitrogen and hydrogen in NH_3 tell us about the composition of the ammonia molecule?

   c. Note that the sum of the reactant coefficients does not equal the sum of the product coefficients for reaction #4. Yet in reaction #2, the sums are equal. Explain why the sums of coefficients do not necessarily have to equal one another in a reaction.

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Chemistry - Unit 6 Reaction Equations Worksheet 1

Balance the following equations by inserting the proper coefficients. For selected reactions, draw *Before* and *After* particle diagrams to show the particles involved in the reaction. Be sure to provide a key.

1. ____C + ____H_2O → ____CO + ____H_2
2. ____MgO → ____Mg + ____O_2
3. ____Al + ____O_2 → ____Al_2O_3

#3 Before After

4. ____Zn + ____H_2SO_4 → ____$ZnSO_4$ + ____H_2
5. ____Cl_2 + ____KI → ____KCl + ____I_2
6. ____CuCl → ____Cu + ____Cl_2
7. ____Na + ____Cl_2 → ____NaCl
8. ____Al + ____HCl → ____$AlCl_3$ + ____H_2

#8 Before After

9. ____Fe_2O_3 → ____Fe + ____O_2
10. ____P + ____O_2 → ____P_2O_5
11. ____Mg + ____HCl → ____$MgCl_2$ + ____H_2
12. ____H_2 + ____N_2 → ____NH_3
13. ____$BaCl_2$ + ____H_2SO_4 → ____$BaSO_4$ + ____HCl
14. ____CH_4 + ____O_2 → ____CO_2 + ____H_2O

#14 Before After

15. a) ____$ZnCl_2$ + ____$(NH_4)_2S$ → ____ZnS + ____ NH_4Cl

b) Find the molar mass of these reactants.

c) How many moles of $ZnCl_2$ would be in 25 g? How much mass would 0.55 moles of $(NH_4)_2S$ have?

Part II: Write the formulas of the reactants and products, then balance the equations. (See Clues and Hints below.)

1. Nitric oxide (NO) reacts with ozone (O_3) to produce nitrogen dioxide and oxygen gas.

2. Iron burns in air to form a black solid, Fe_3O_4.

3. Sodium metal reacts with chlorine gas to form sodium chloride.

4. Acetylene, C_2H_2, burns in air to form carbon dioxide and water.

5. Hydrogen peroxide (H_2O_2) easily decomposes into water and oxygen gas.

6. Hydrazine (N_2H_4) and hydrogen peroxide are used together as rocket fuel. The products are nitrogen gas and water.

7. If potassium chlorate is strongly heated, it decomposes to yield oxygen gas and potassium chloride.

8. When sodium hydroxide is added to sulfuric acid (H_2SO_4), the products are water and sodium sulfate.

9. In the Haber process, hydrogen gas and nitrogen gas react to form ammonia, NH_3.

CLUES and HINTS:

- Products usually follow words like *produces, yields, forms*
- Watch for our diatomic elements (*H_2,N_2, etc…)*, which are often (but not always) <u>gases</u>
- Include 'state subscripts' behind each substance [(s), (l), (g)] when the state is given
- Remember **air** is a mixture of (primarily) two gases, O_2 and N_2. Which is most likely to participate in a reaction?
- Elemental metals exist as single, unbonded atoms. (Ex: formula for copper metal is **Cu**)
- Watch for **ionic** vs **molecular** compounds. Use *nomenclature rules,* and your *ion chart* and *periodic table* to figure out the formulas for these.

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Unit 6: Balancing Chemical Reactions- Worksheet 2

Balance the following equations by inserting the proper coefficients.

1. ____SO_2 + ____O_2 → ____SO_3
2. ____CH_4 + ____O_2 → ____CO + ____H_2O
3. ____P + ____Cl_2 → ____PCl_3
4. ____CO + ____O_2 → ____CO_2
5. ____CH_4 + ____ O_2 → ____CH_3OH
6. ____Li + ____Br_2 → ____$LiBr$
7. ____Al_2O_3 → ____Al + ____O_2
8. ____Na + ____H_2O → ____$NaOH$ + ____H_2
9. ____CO_2 + ____H_2O → ____$C_6H_{12}O_6$ + ____O_2
10. ____H_2SO_4 + ____$NaCl$ → ____HCl + ____Na_2SO_4

Two reactions used to get rid of sulfur dioxide, a pollutant from burning coal:

11. ____H_2 + ____SO_2 → ____H_2S + ____H_2O
12. ____$CaCO_3$ + ____SO_2 + ____ O_2 → ____$CaSO_4$ + ____CO_2

13. ____$AgNO_3$ + ____$CaCl_2$ → ____$AgCl$ + ____$Ca(NO_3)_2$
14. ____HCl + ____$Ba(OH)_2$ → ____$BaCl_2$ + ____H_2O
15. ____H_3PO_4 + ____$NaOH$ → ____Na_3PO_4 + ____H_2O
16. ____$Pb(NO_3)_2$ + ____KI → ____ PbI_2 + ____KNO_3
17. ____CuO + ____NH_3 → ____N_2 + ____Cu + ____H_2O
18. ____C_2H_5OH + ____O_2 → ____CO_2 + ____H_2O
19. ____C_2H_6 + ____O_2 → ____CH_3COOH + ____H_2O
20. ____NO_2 + ____H_2O → ____HNO_3 + ____NO

Write the formulas of the reactants and products - including the symbols for the state - (s), (l), (g), (aq) - then balance the equations.

1. When a solution of hydrogen chloride is added to solid sodium bicarbonate ($NaHCO_3$), the products are carbon dioxide, water and aqueous sodium chloride.

2. Steam (gaseous water) reacts with carbon at high temperatures to produces carbon monoxide and hydrogen gases.

3. Limestone, $CaCO_3$, decomposes when heated to produce lime, CaO, and gaseous carbon dioxide.

4. Ethyl alcohol (a liquid), C_2H_6O, burns in air to produce carbon dioxide and gaseous water.

5. Solid titanium(IV) chloride reacts with water, forming solid titanium(IV) oxide and aqueous hydrogen chloride.

6. At high temperatures, the gases chlorine and water react to produce hydrogen chloride and oxygen gases.

7. Steel wool (nearly pure Fe) burns in air to form the solid iron oxide, Fe_2O_3.

8. During photosynthesis in plants, carbon dioxide and water are converted into glucose, $C_6H_{12}O_6$, and oxygen gas.

9. Solutions of calcium hydroxide, $Ca(OH)_2$and nitric acid, HNO_3, react to produce water and aqueous calcium nitrate, $Ca(NO_3)_2$.

Name ______________________________

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Chemistry – Unit 6 Worksheet 3

Write balanced chemical equations for the following reactions.

1. Ammonia (NH_3) reacts with hydrogen chloride to form ammonium chloride.

2. Calcium carbonate decomposes upon heating to form calcium oxide and carbon dioxide.

3. Barium oxide reacts with water to form barium hydroxide.

4. Acetaldehyde (CH_3CHO) decomposes to form methane (CH_4) and carbon monoxide.

5. Zinc reacts with copper(II) nitrate to form zinc nitrate and copper.

6. Calcium sulfite decomposes when heated to form calcium oxide and sulfur dioxide.

7. Iron reacts with sulfuric acid (H_2SO_4) to form iron(II) sulfate and hydrogen gas.

8. A nitrogen containing carbon compound, $C_2H_6N_2$, decomposes to form ethane, C_2H_6, and nitrogen gas.

9. Phosgene, $COCl_2$, is formed when carbon monoxide reacts with chlorine gas.

10. Manganese(II) iodide decomposes when exposed to light to form manganese and iodine.

11. Dinitrogen pentoxide reacts with water to produce nitric acid (HNO_3).

12. Magnesium reacts with titanium (IV) chloride to produce magnesium chloride and titanium.

13. Carbon reacts with zinc oxide to produce zinc and carbon dioxide.

14. Bromine reacts with sodium iodide to form sodium bromide and iodine.

15. Phosphorus (P_4) reacts with bromine to produce phosphorus tribromide.

16. Ethanol, C_2H_5OH, reacts with oxygen gas to produce carbon dioxide and water.

17. Calcium hydride reacts with water to produce calcium hydroxide and hydrogen gas.

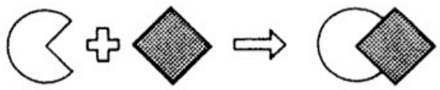
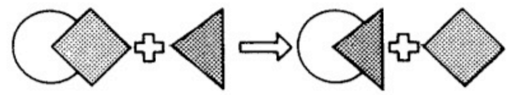
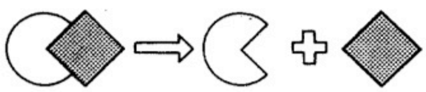
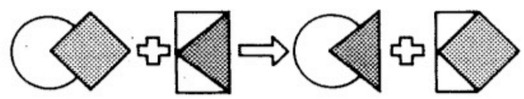
18. Sulfuric acid, H_2SO_4, reacts with potassium hydroxide to produce potassium sulfate and water.

19. Propane, C_3H_8, burns in air to produce carbon dioxide and water.

NAME: ____________________

Identifying Reactions and Balancing Equations

The sample equations below show the four types of reactions and an example of each.

Synthesis Reaction **A + B → AB**  $C + O_2 \rightarrow CO_2$	**Single Displacement Reaction** **A + BC → B + AC**  $Zn + 2HCl \rightarrow H_2 + ZnCl_2$
Decomposition Reaction **AB → A + B**  $2H_2O \rightarrow 2H_2 + O_2$	**Double Displacement Reaction** **AB + CD → AD + CB**  $HCl + NaOH \rightarrow H_2O + NaCl$

Identify the type of reaction for items 1–12 below. Balance any equations that are not balanced. REMEMBER: When balancing a chemical equation, only the coefficients change. The subscripts and chemical formulas do not change.

Place the proper coefficient in the blanks provided. — **Reaction Type**

1. _____ KNO_3 → _____ KNO_2 + _____ O_2 ____________________
2. _____ $AgNO_3$ + _____ $ZnCl_2$ → _____ $AgCl$ + _____ $Zn(NO_3)_2$ ____________________
3. _____ $CdSO_4$ + _____ H_2S → _____ CdS + _____ H_2SO_4 ____________________
4. _____ Al + _____ HCl → _____ $AlCl_3$ + _____ H_2 ____________________
5. _____ C + _____ $BaSO_4$ → _____ CO + _____ BaS ____________________
6. _____ $Zn(OH)_2$ → _____ ZnO + _____ H_2O ____________________
7. _____ Fe_3O_4 + _____ H_2 → _____ Fe + _____ H_2O ____________________
8. _____ CaS + _____ HCl → _____ $CaCl_2$ + _____ H_2S ____________________
9. _____ Cl_2 + _____ $NaBr$ → _____ $NaCl$ + _____ Br_2 ____________________
10. _____ SiF_4 + _____ H_2O → _____ HF + _____ $Si(OH)_4$ ____________________
11. _____ HCl + _____ Fe_2O_3 → _____ $FeCl_3$ + _____ H_2O ____________________
12. _____ $BaCO_3$ → _____ Ba + _____ CO_2 ____________________

Name ______________________ Date ____________ Class ____________

CHAPTER 9 REVIEW ACTIVITY

Text Reference: **Section 9-12**

Categories of Chemical Reactions

State whether each of the following equations represents a synthesis (s), analysis (a), single replacement (sr), or double replacement (dr) reaction.

1. $CO_2 \rightarrow C + O_2$ 1. ____________
2. $NaCl + AgNO_3 \rightarrow NaNO_3 + AgCl$ 2. ____________
3. $S + Cl_2 \rightarrow SCl_2$ 3. ____________
4. $BaCl_2 + 2NaOH \rightarrow 2NaCl + Ba(OH)_2$ 4. ____________
5. $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$ 5. ____________
6. $CH_4 \rightarrow C + 2H_2$ 6. ____________
7. $Pb(NO_3)_2 + Mg \rightarrow Pb + Mg(NO_3)_2$ 7. ____________
8. $Mg + 2HCl \rightarrow MgCl_2 + H_2$ 8. ____________
9. $H_2SO_4 \rightarrow H_2 + S + 2O_2$ 9. ____________
10. $2O_2 + N_2 \rightarrow N_2O_4$ 10. ____________
11. $3CaBr_2 + 2Na_3P \rightarrow Ca_3P_2 + 6NaBr$ 11. ____________
12. $2KI + Br_2 \rightarrow 2KBr + I_2$ 12. ____________
13. $C_6H_{12}O_6 \rightarrow 6C + 6H_2O$ 13. ____________
14. $2NaF \rightarrow 2Na + F_2$ 14. ____________
15. $Si + O_2 \rightarrow SiO_2$ 15. ____________
16. $2NaI + Pb(NO_3)_2 \rightarrow 2NaNO_3 + PbI_2$ 16. ____________
17. $NaI + Cs \rightarrow CsI + Na$ 17. ____________
18. $H_2 + CO + O_2 \rightarrow H_2CO_3$ 18. ____________
19. $Li_3PO_4 \rightarrow 3Li + P + 2O_2$ 19. ____________
20. $CS_2 + 2F_2 \rightarrow CF_4 + 2S$ 20. ____________

Describing Chemical Reactions

Introduction and Purpose

In this experiment you will observe examples of the five basic types of chemical reactions. You will learn to write balanced equations, including the role of energy, to effectively communicate the chemistry of the reactions.

Procedure

Carry out the reactions using the approximate quantities of reagents suggested. Unless otherwise stated, use test tubes. When heating reagents in test tubes, slant the test tube so that the opening is pointed *away* from people. Heat the test tube at the surface of the material and work down towards the bottom of the tube. Discard solutions down the drain, wash and rinse your glassware. Discard solid waste in the waste cans on the lab tables.

In the data section you will balance the equation, write the word equation and record your observations.

A. Combination reactions:

1. Grasp a strip of magnesium ribbon in crucible tongs and ignite it in the burner flame. Hold it over a watch glass. Do not look directly at the flame! Add a few drops of distilled H_2O to the ash. Stir with a stirring rod and place a drop of the solution on red litmus paper. Red litmus turning blue is evidence for the presence of a base.

2. Heat a piece of copper metal strongly in the Bunsen burner flame for about 30 s. Remove the copper from the flame and note the change in appearance. Discard the product in the solid waste can.

B. Decomposition reactions:

1. Place about 1 scoopful of solid sodium hydrogen carbonate $NaHCO_3$ into a dry test tube. Mass the test tube with the powder. Heat the sodium hydrogen carbonate in the test tube strongly for 2 minutes. Observe any changes that occur during the heating. Toward the end of the heating, light a wood splint and insert the flaming splint into the mouth of the test tube. Note what happens to the splint. Once the tube has cooled, mass the tube and contents again.

C. Single replacement reactions:

1. Place a strip of copper in a test tube with enough 0.1M $AgNO_3$ to cover it. Set this test tube aside, then observe the surface of the metal after 5-10 minutes.

2. Place a couple of pieces of mossy zinc metal in a test tube approximately 1/4 full of 3M HCl. Place a stopper loosely in the tube. After a few minutes, light a wood splint and insert the flaming splint into the mouth of the test tube. Hold the test tube in your hand to feel if the temperature has changed.

D. Double replacement reactions:

1. Add 0.1M $AgNO_3$ to a test tube to a depth of about 1 cm. Add a similar quantity of 0.1M $CaCl_2$ solution. Observe the reaction.

2. Place a scoopful of solid Na_2CO_3 in a test tube to a depth of about 1 cm. Add a dropperful of 3M HCl. While the reaction is occurring, test with a flaming splint as in part B.
   Check to see if the temperature of the mixture has changed.

E. Combustion reactions:

Place about 10 drops of isopropyl alcohol, C_3H_7OH, in a small evaporating dish. Ignite the alcohol from the top of the liquid with a Bunsen burner. Hold a cold watch glass well above the flame and observe the condensation of water on the bottom. The formation of the mist will be fleeting; watch closely.

Lab Report

Your lab report should include the **purpose of the lab**, the **completed data and evaluation sheet**, and answers (in complete sentences) to the **following questions**.

1. What are some of the observable changes that are evidence that a chemical reaction has taken place?

2. How did the flaming splint behave when it was inserted into the tube with $CO_{2\,(g)}$? In what way was this different from the reaction of the $H_{2(g)}$ to the flaming splint?

3. In the reaction of magnesium with oxygen gas, a considerable amount of energy was released. This is an example of an *exothermic* reaction. From this evidence what can you conclude about the energy stored in the reactants compared to the energy stored in the product? What other examples of exothermic reactions did you observe? Re-write the balanced equation for the reaction of Mg and O_2, this time with the term "+ energy" on the appropriate side of the equation.

4. You had to heat the $NaHCO_3$ strongly in order for it to decompose. This is an example of an *endothermic* reaction. What does this tell you about the energy stored in the reactants compared to the energy stored in the product? Write the balanced equation for the decomposition of $NaHCO_3$, this time with the term "+ energy" on the appropriate side of the equation.

Data and Evaluation

Record your observations and balance the equations in the section below.

A. Combination reactions:

1a. Observations:

1a. _____ Mg + ____O_2 → _____ MgO

Write in words:

1b. Observations:

1b. ______ MgO + __ H_2O → ___ $Mg(OH)_2$ (aq)

Write in words:

2. Observations:

2. ______ Cu + _____ O_2 --> _____CuO

Write in words:

B. Decomposition reactions

1. Observations:

1. ______ $NaHCO_3$ → ____ Na_2O + ______ H_2O + _____ CO_2 (g)

Write in words:

C. Single replacement reactions

1. Observations:

1. ______ $AgNO_3$ (aq) + __ Cu → _____ Ag + _____ $Cu(NO_3)_2$ (aq)

Write in words:

2. Observations

2. ______ Zn + _____ $HCl_{(aq)}$ → _______$ZnCl_{2(aq)}$ + ______$H_{2(g)}$
Write in words:

D. Double replacement reactions

1. Observations

1. ______$AgNO_{3(aq)}$ + _____ $CaCl_{2(aq)}$ → _____ $AgCl_{(s)}$ + _______ $Ca(NO_3)_{2(aq)}$
Write in words:

2. Observations

2. ______ Na_2CO_3 + ______ HCl $_{(aq)}$ → ____________$NaCl_{(aq)}$ + ______ H_2O + ___$CO_{2(g)}$
Write in words:

E. Combustion reactions

1. Observations

1. ______$C_3H_7OH_{(l)}$ + _____ $O_{2(g)}$ → ___ $CO_{2(g)}$ + __$H_2O_{(g)}$
Write in words:

Patterns of Chemical Reaction

Just as there are patterns in the way elements bond and the types of compounds they make, there are patterns in the way chemical bonds are rearranged in a reaction. These patterns make it easier to predict the outcome of a chemical reaction.

Below are 5 useful reaction patterns. Look at the examples given for each type. Clues for identifying these patterns include the kind of substances involved (ionic or molecular), whether certain elements (like O_2) are always present, or if the complexity of the molecules goes up or down.

1. **Combination** or **Synthesis** ('building up'– fewer **product** molecules or more complex **product** molecules)
   $2\ Na + Cl_2 \rightarrow 2\ \mathbf{NaCl}$
   $6\ CO_2 + 6\ H_2O \rightarrow \mathbf{C_6H_{12}O_6} + 6\ O_2$ (this one is called photo*synthesis*!)
   $2\ Li + Br_2 \rightarrow 2\ \mathbf{LiBr}$

2. **Decomposition** ('breaking down'– more complex **reactant** molecules form simpler **product** molecules)
   $2\ \mathbf{NaCl} \rightarrow 2\ Na + Cl_2$
   $\mathbf{C_6H_{12}O_6} + 6\ O_2 \rightarrow 6\ CO_2 + 6\ H_2O$
   $2\ \mathbf{LiBr} \rightarrow 2\ Li + Br_2$

*NOTE: What is the relationship between **synthesis** and **decomposition** reactions?*

3. **Replacement** (or *single replacement* – one element replaces another in a compound)
   $2\ \mathbf{Fe} + 3\ \mathbf{Cu}Cl_2 \rightarrow 3\ \mathbf{Cu} + 2\ \mathbf{Fe}Cl_3$ (Fe replaces Cu: is this familiar?)
   $2\ \mathbf{Na} + 2\ \mathbf{H}OH$ (*or H_2O*) $\rightarrow \mathbf{H_2} + 2\ \mathbf{Na}OH$ (Na replaces H – metal replacement)
   $\mathbf{Ni} + 2\ \mathbf{H}Cl \rightarrow \mathbf{H_2} + \mathbf{Ni}Cl_2$ (Ni replaces H – metal replacement)
   $2\ \mathbf{F_2} + 2\ H_2\mathbf{O} \rightarrow \mathbf{O_2} + 4\ H\mathbf{F}$ (F replaces O – nonmetal replacement)

4. **Double Replacement** (between aqueous ionic compounds: ions trade partners; one product is insoluble in water and *precipitates*)
   $\mathbf{Ca}Cl_2(aq) + 2\ \mathbf{Ag}NO_3(aq) \rightarrow \mathbf{Ca}(NO_3)_2(aq) + 2\ \mathbf{Ag}Cl(s)$ (Ca, Ag trade; AgCl precipitates)
   $\mathbf{Na_2}SO_4(aq) + \mathbf{Ba}Cl_2(aq) \rightarrow 2\ \mathbf{Na}Cl(aq) + \mathbf{Ba}SO_4(s)$ (Na, Ba trade; $BaSO_4$ precipitates)
   $\mathbf{Pb}(NO_3)_2(aq) + 2\ \mathbf{Li}OH(aq) \rightarrow \mathbf{Pb}(OH)_2(s) + 2\ \mathbf{Li}NO_3(aq)$ (Pb, Li trade; $Pb(OH)_2$ precipitates)

5. **Combustion** (combines with oxygen; most or all products contain oxygen; releases energy to environment)

$CH_4 + 2\ \mathbf{O_2} \rightarrow C\mathbf{O_2} + 2\ H_2\mathbf{O}$ + energy (burning natural gas)

$C_3H_8 + 5\ \mathbf{O_2} \rightarrow 3\ C\mathbf{O_2} + 4\ H_2\mathbf{O}$ + energy (burning propane)

$2\ Mg + \mathbf{O_2} \rightarrow 2\ Mg\mathbf{O}$ + energy (can be classified as *combustion* or *synthesis*)

$4\ Fe + 3\ \mathbf{O_2} \rightarrow 2\ Fe_2\mathbf{O_3}$ + energy (can be classified as *combustion* and *synthesis*)

ASSIGNMENT:

Examine the reactions from Balancing Equations ws 2. Identify which of these patterns applies to each reaction on both pages (except p1, #11, 12, 17, and 20, which don't fit the patterns above). Write the pattern name in the margin next to each reaction.

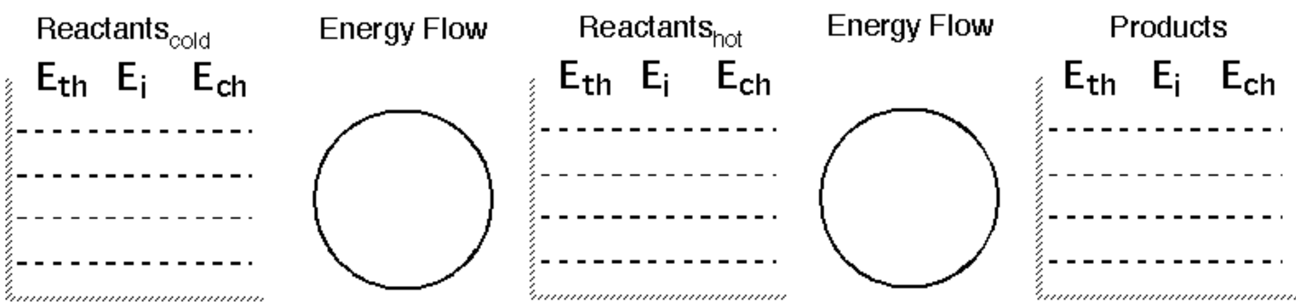
Name ____________________

Date ______________ Pd ___

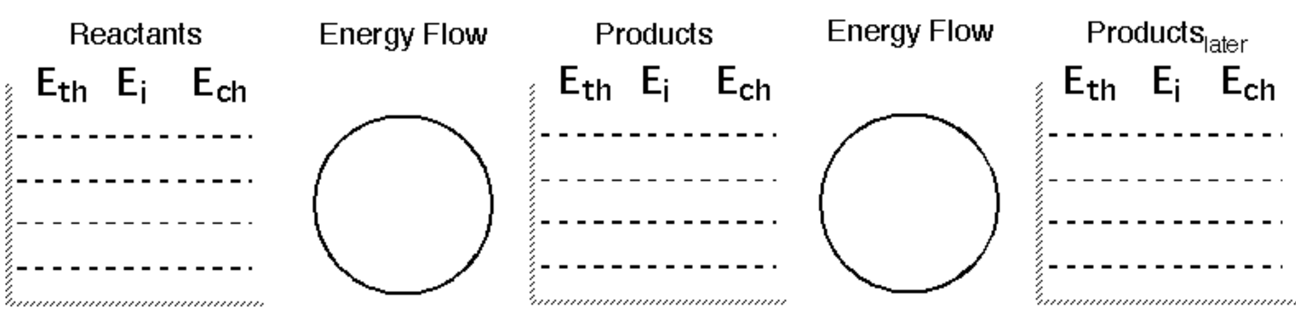
Chemistry – Unit 6 Worksheet 4
Representing Chemical Potential Energy in Change

For each of the reactions below, write the balanced chemical equation, including the energy term on the correct side of the equation. Then represent the energy storage and transfer using the bar graphs. Below the bar graph diagram, sketch a standard chemical potential energy curve for the reaction.

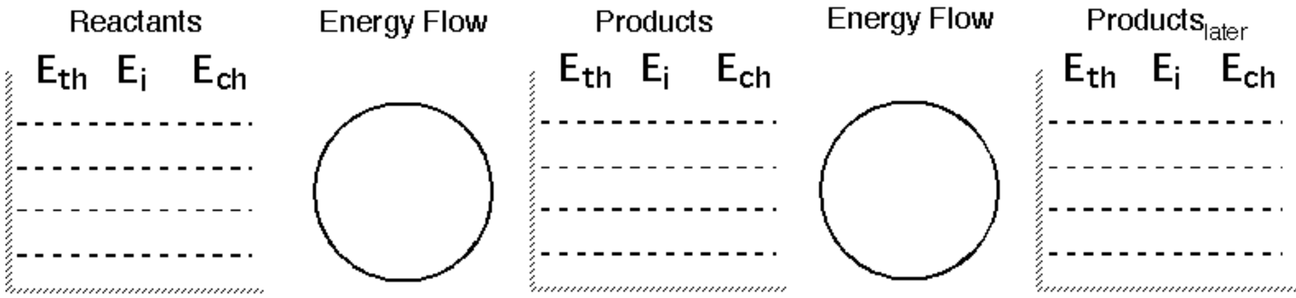
1. When you heated sodium hydrogen carbonate, you decomposed it into sodium oxide, water vapor, and gaseous carbon dioxide.



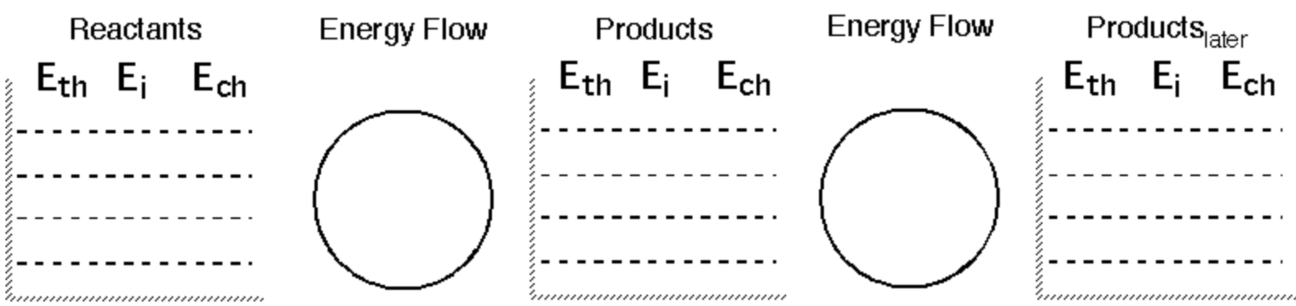
2. When solid zinc was added to hydrochloric acid, the products were hydrogen gas and an aqueous solution of zinc chloride. You could feel the test tube get hotter.



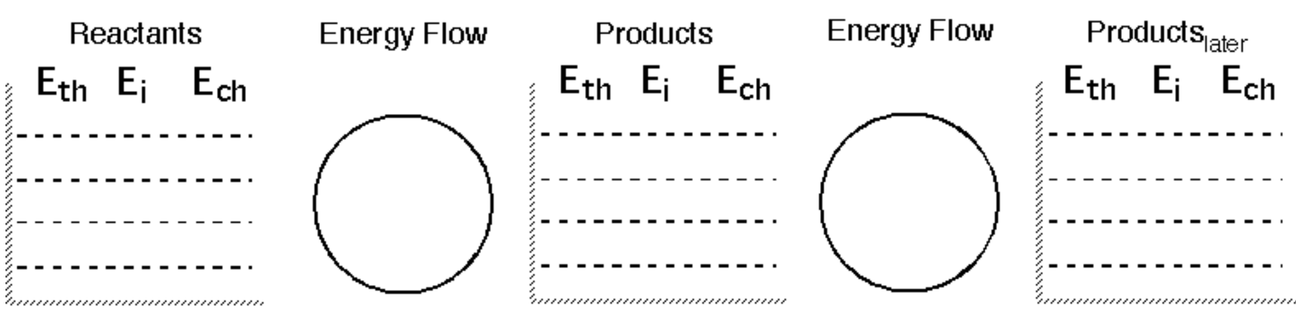
3. Isopropyl alcohol burned in air to produce carbon dioxide and water vapor.



4. In chemical cold packs, solid ammonium chloride dissolves in water forming aqueous ammonium and chloride ions. As a result of this solvation reaction, the pack feel cold on your injured ankle.



5. In chemical hot packs, solid sodium acetate crystallizes from a supersaturated solution of sodium acetate. The pack feels warm to the touch for 30 minutes or longer.



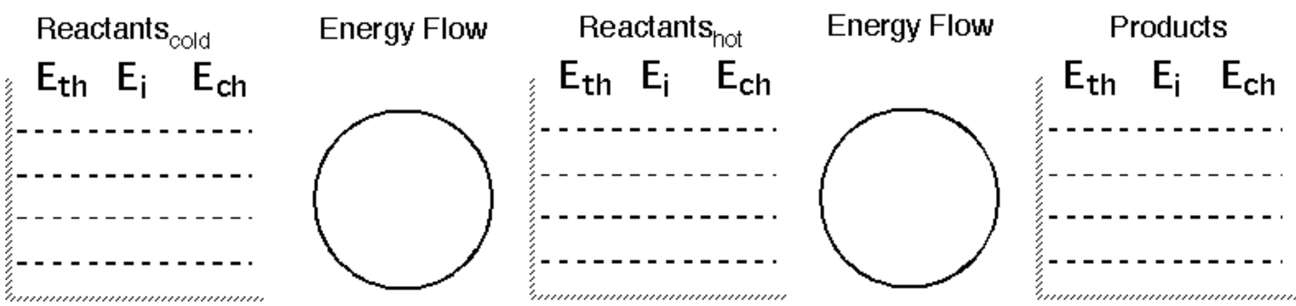
Name ___________________________________

Date ________________ Pd ___

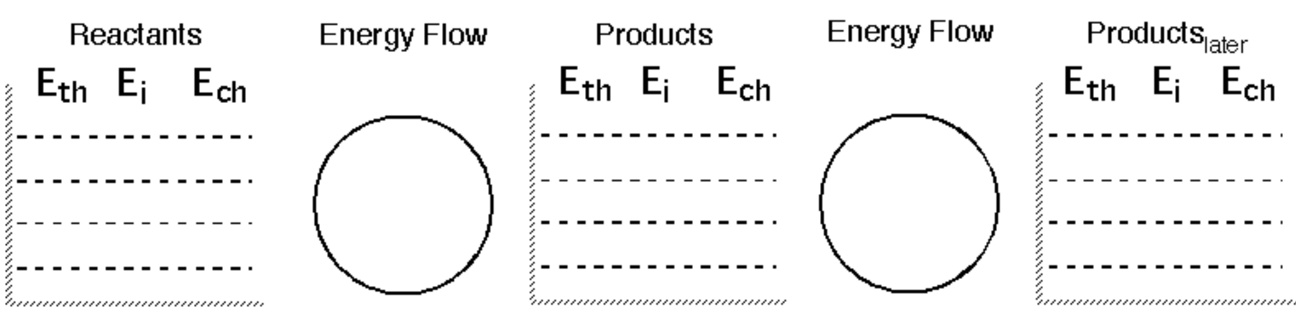
Chemistry – Unit 6 Worksheet 4A
Representing Chemical Potential Energy in Change

For each of the reactions below, write the balanced chemical equation, including the energy term on the correct side of the equation. Then represent the energy storage and transfer using the bar graphs. Below the bar graph diagram, sketch a standard chemical potential energy curve for the reaction.

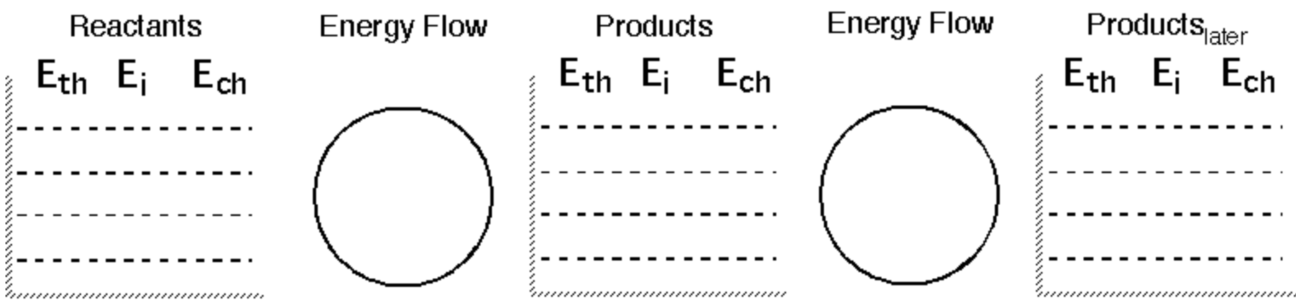
1. You crack an egg and place it in a hot frying pan. In a few seconds, the egg white turns from clear liquid to opaque white solid.



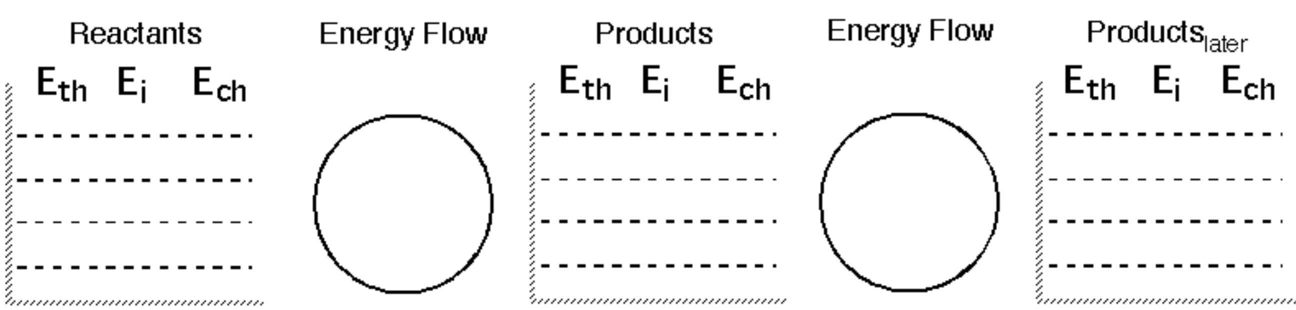
2. Potassium metal is dropped into a flask of water. The reaction of the metal with the water gets hot, flames shoot up and hydrogen gas as well as aqueous potassium hydroxide are formed.



3. Vinegar (hydrogen acetate) is mixed into sodium bicarbonate. The resulting "volcano" reaction forms lots of bubbles, water and sodium acetate. The container also feels cooler after the reaction.



4. Iron powder is left in a beaker exposed to the air. A thermometer in the beaker shows that the beaker gets warmer over time. The next day, most of the iron has formed iron (III) oxide.



6. Chunks of sodium hydroxide are dropped into a flask of water. During the dissolving, the flask gets very hot. Aqueous sodium hydroxide is formed.